Bell Work
9-Dec-13
On New BW #1x

How are group number *and* number of valence electrons related?

What is Lewis dot structure of an atom?

Agenda

Lewis Dot Structure – Atoms (recap),
Compounds
Steps 1-4

Objective:
You will KNOW how to draw Lewis Structures of simple common compounds and ions

Turn In
9-Dec-2013

Precipitation Lab
BW #10

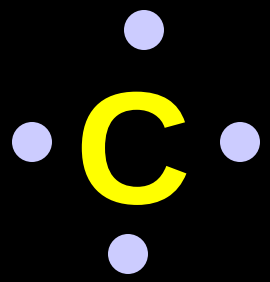
Drawing Lewis Structures

We learned how to draw Lewis Dot structures of ions a few months ago.

To draw a Lewis dot structure of an ion simply draw the elements symbol and then distribute electrons around the symbol.

Take Carbon:

Total # of valence e^- (equals group #): 4



You try: F, O, Mg

You have 30 sec.

Drawing Lewis Structures

Lewis structures are used to identify the types of bonds (single —, double ═, triple ≡) formed between atoms in a molecule or polyatomic ion.

Drawing the Lewis structure is not difficult ☺
IF you follow the exact process that I give you.

Now for the steps…

Drawing Lewis Structures

1. Add up the valence electrons from all atoms

H_2O Ve^- :
1 from each H & 6 from O
So 2(1) + 1(6) =8

CO_2 Ve^-:
4 from C & 6 from each O
So 1(4) + 2(6) = 16

You try: SO_2 and SiO_2
SO_2 Ve^- = 1(6) + 2(6) = 18

SiO_2 Ve^- = 1(4) + 2(6) = 16

Drawing Lewis Structures

1. Add up the valence electrons from all atoms

For a cation (+), subtract 1 electron for each positive charge

NH_4^+ : 1(5) + 4 (1) -1 = 8 Ve^-

For an anion (-), add 1 electron for each negative charge

CN^- : 1(4) + 1(5) + 1 = 10 Ve^-

You try: NO_2^-, CO_3^{2+}

NO_2^- Ve^- = 1(5) + 2(6) + 1 = 18

CO_3^{2+} Ve^- = 1(4) + 3(6) -2 = 20

Drawing Lewis Structures

2. Draw a skeleton structure showing the chemical symbols for each atom. Connect the appropriate atoms using a single bond –, each line represent 2 e-.

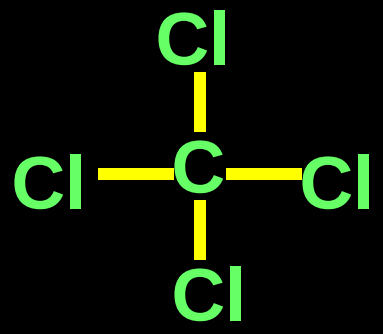
Sometimes (but not always) the order in which the formula is written

HCN: H-C-N

Central atom (written first) surrou

atoms

CCl_4 :



Drawing Lewis Structures

3. Add electron pairs, ••,to the atoms bonded to the central atom first until each has an octet (8) of e^-.

Remember, H only gets $2e^-$ so once it bonds it has its $2e^-$.

H — Cl

IF there are any unused e^-, *place all of the leftovers on the central atom*.

HONORS

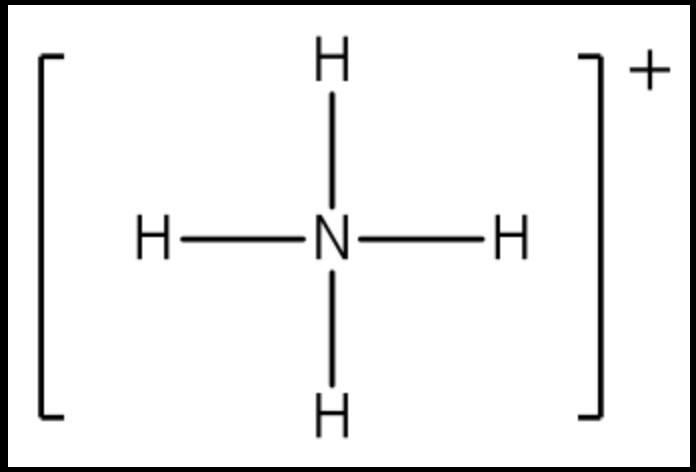
Note: This sometimes gives the central atom more than eight e^-.

Drawing Lewis Structures

4. Do all atoms that need an octet have one?
Did you use all of the valance electrons?

If you answered yes then you are done.

Note: if you are drawing an ion (charged particle) the you must put the structure in brackets and label the charge NH_4^+



Practice

Complete the following for Br_2

#of Valence electrons :____

# of lone pairs (electrons ••):_____

#of bonding pairs (—): _____

Structure:

Br Br

Drawing Lewis Structures

- **Lets try one: H_2O**

Number of Ve^- : **2(1) + 1(6) = 8e^-**

-4e^-

-4e^-

0

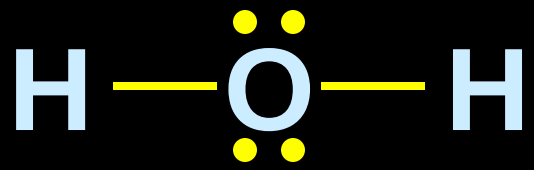
Did we use all the Ve^-?

Do all the atoms that need an octet have one?

Yesss, you have done good job!!!

Drawing Lewis Structures

H_2O



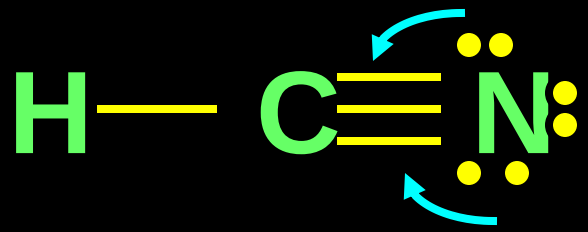
We have
2 lone e- pairs
2 bonding pairs

Drawing Lewis Structures

- If **(and only if)** there are not enough e^- to give the central atom an octet, try multiple bonds.

Use one (or more) unshared pairs of e^- to form double (or triple) bonds: **HCN**

Number of Ve^- : 1+ 4 + 5 = 10Ve^-



-4e^-

-6e^-

-0Ve^-

Oops Carbon does not have an octet…

Now both nitrogen and carbon have an octet

Drawing Lewis Structures

Example: Draw the Lewis structure for $CHCl_3$

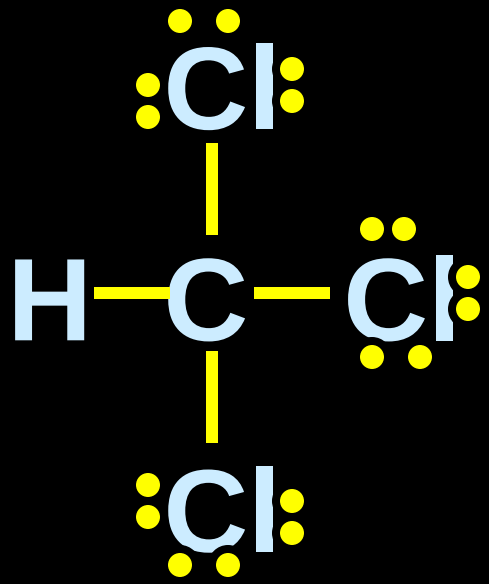
# of Ve^- = 4 + 1 + 3(7) = 26Ve^-

C = central atom

-8e^-

-18e^-

0Ve^-



We have used all the Ve- and every atom that needs and octet has one

Drawing Lewis Structures

Example: Draw the Lewis structure for PO_4^{3-}.

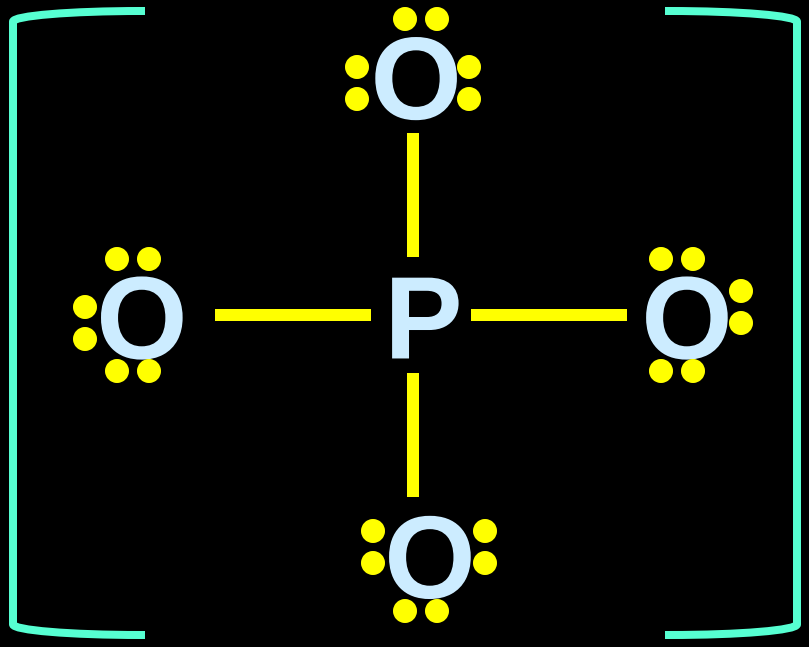
 # valence electrons = 5 + 4(6) + 3 = 32Ve$^-$

P = central atom

-8e$^-$

-24e$^-$

-0Ve$^-$



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Don't forget to show the charge of the ion, too.

Drawing Lewis Structure

With your partner write the steps for drawing Lewis structures. Note any special manipulations you may need to do to get all atoms an octet and use all your Ve^-.

1.

2.

3.

4.

Small Group Practice

In your lab groups please complete the Lewis structures of the following:

Carbon dioxide

Elemental iodine

CH_3Cl

Sulfate ion* **(remember the charge)**

Bell Work
10-Dec-2014

What amount of energy is produce when a 150nm wavelength of light is emitted?

Get a text book when finished

Objective:

You will KNOW how to predict the molecular geometry and bond angles of simple compounds based on their Lewis Structure

Practice

Complete the following for SiO
#of Valence electrons :____
# of lone pairs (electrons ••):_____
#of bonding pairs (‒): _____
Structure:

Si O

Drawing Lewis Structures
Your Turn

Draw the Lewis structure for PCl_3

Draw the Lewis structure for NO_2^-

Draw the Lewis structure for XeF_2

Drawing Lewis Structures
Ion practice

a) BH_2^-

b) SO_3^{2-}

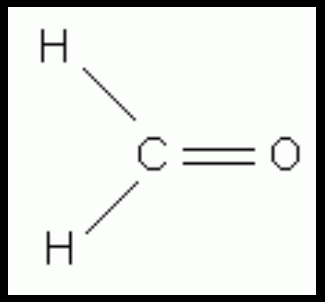
c) ClF_4^+ (Cl gets more than an octet)

d) ClO_3^-

Drawing Lewis Structures
Ion practice

Which is the correct structure for CH_2O?

1



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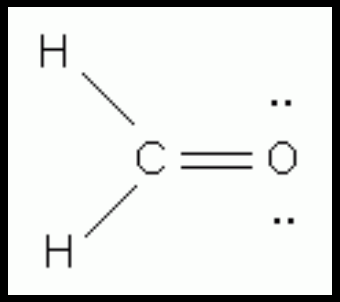

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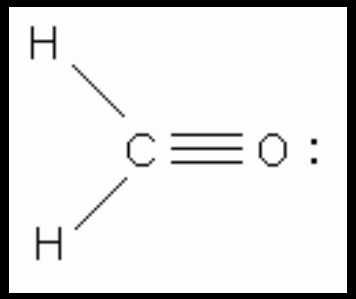

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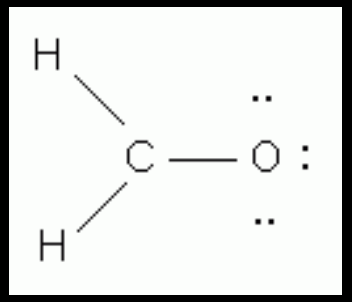


Drawing Lewis Structures

When writing the Lewis structure for ozone we could easily have put the double bond between the other two oxygens.

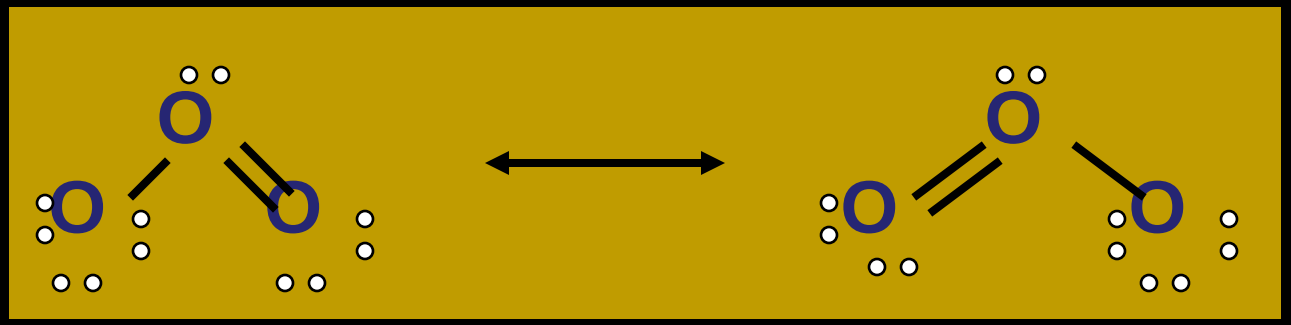
Ozone (O_3):

Drawing Lewis Structures

These two structures are equivalent except for the placement of electrons.

Resonance structures

Resonance structures for ozone:



Drawing Lewis Structures

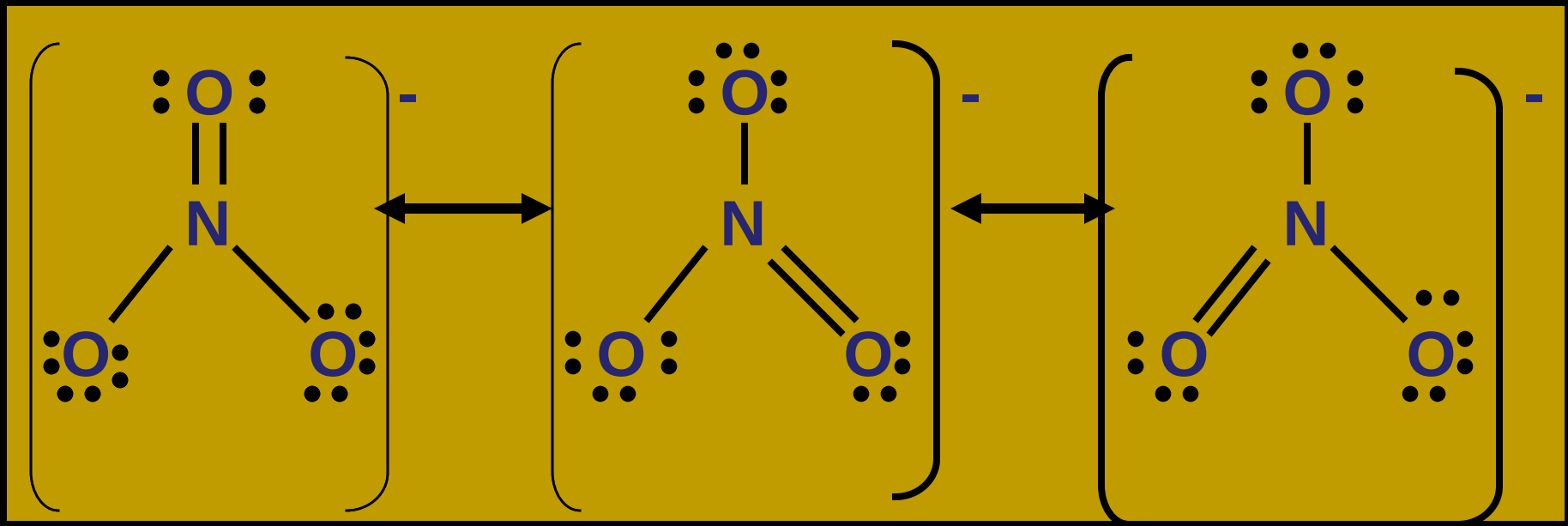
Resonance structure: **one of a group of Lewis structures used to describe a molecule that cannot be accurately depicted using a single Lewis structure**

NOTE: The real molecule is a "hybrid" or average of the resonance structures. It does not "flip" back and forth between the possible structures.

Drawing Lewis Structures

Example: Draw all possible resonance structures for NO_3^-.

 # valence electrons = 5 + 3(6) + 1 = 24



Review of Chemical Bonds

There are 3 forms of bonding:

ionic —complete *transfer* of 1 or more electrons from one atom to another (one loses, the other gains) forming oppositely charged ions that attract one another (metal and non-metal)

Covalent —some valence electrons *shared* between atoms (non-metal and non-metal)

Metallic – holds atoms of a metal together (two or more metals)

The type of bond can *usually* be calculated by finding the difference in electronegativity of the two atoms that are going together.

Page 303 Table 12.1

H 2.1

1A	2A						8B			1B	2B	3A	4A	5A	6A	7A
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	3B	4B	5B	6B	7B						Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2

<1.0 | 1.0–1.4 | 1.5–1.9 | 2.0–2.4 | 2.5–2.9 | 3.0–4.0

Figure 9.9 Electronegativity values for the elements according to Pauling. Trends for electronegativities are the opposite of the trends defining metallic character. Nonmetals have high values of electronegativity, the metalloids have intermediate values, and the metals have low values.

Electronegativity Difference

If the difference in electronegativities is between:

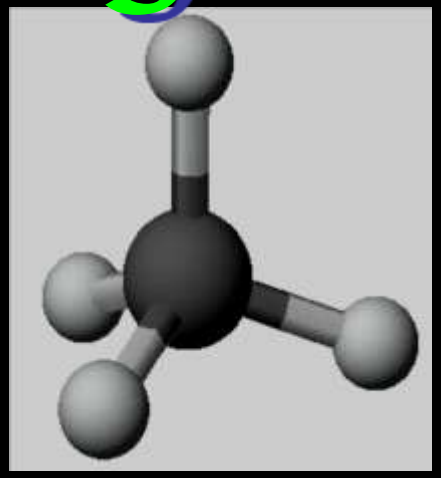
1.7 to 4.0: Ionic

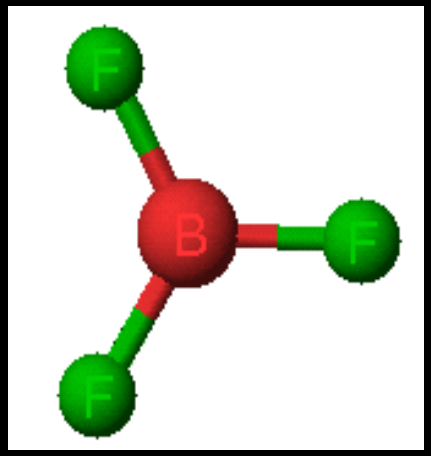
0.3 to 1.7: Polar Covalent

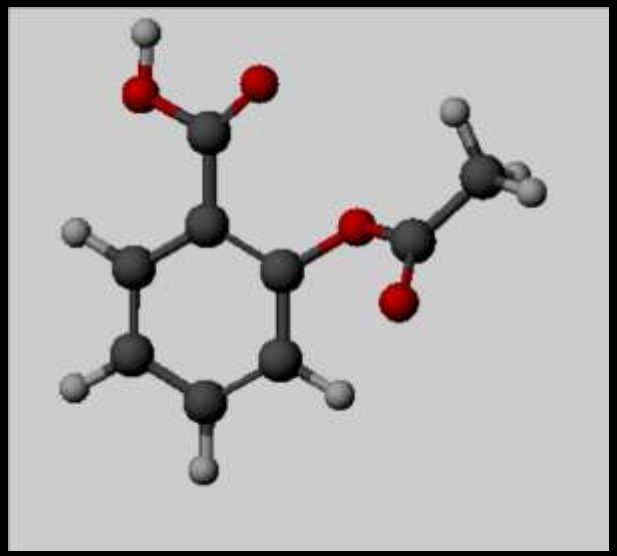
0.0 to 0.3: Non-Polar Covalent

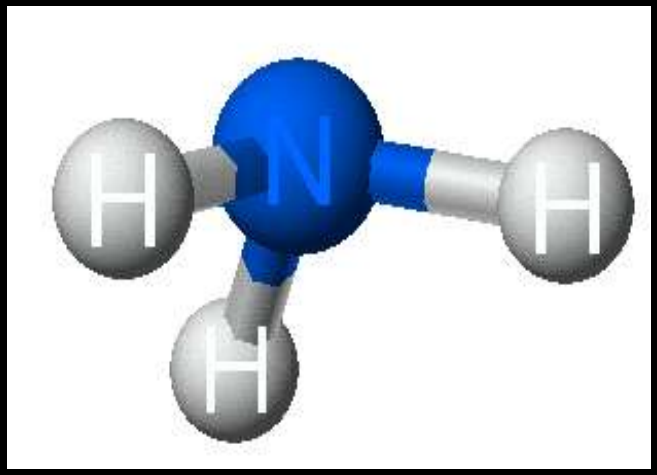
Example: NaCl
Na = 1.01, Cl = 2.83
Difference is 1.82, so this is an ionic bond!

MOLECULAR GEOMETRY









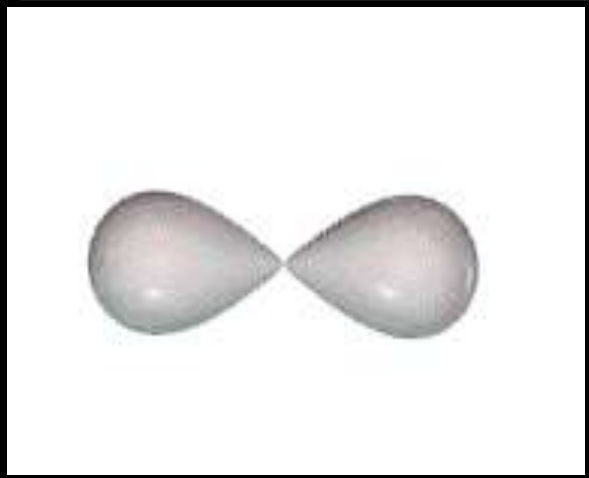
MOLECULAR GEOMETRY

VSEPR

Valence **S**hell **E**lectron **P**air **R**epulsion theory

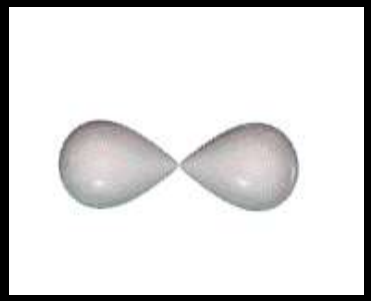
Most important factor in determining geometry is

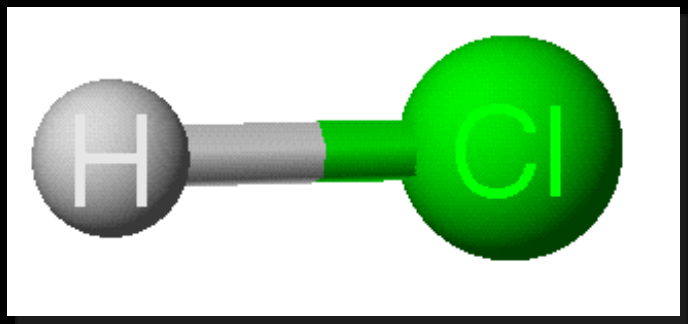
Molecule adopts the shape that minimizes the electron pair repulsions.



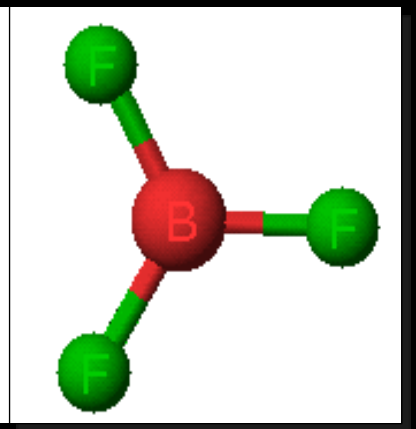
Some Common Geometries

Linear

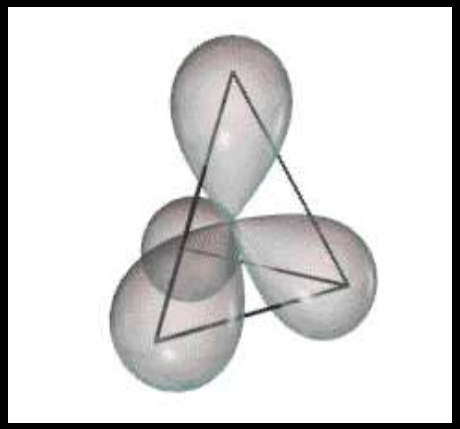


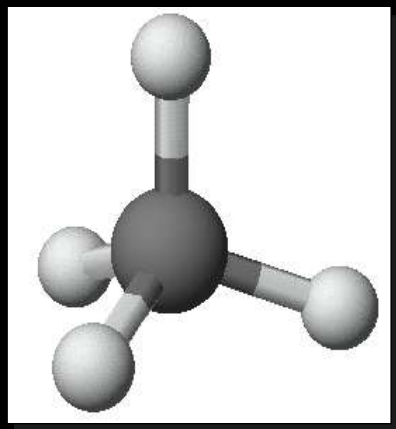


Trigonal Planar



Tetrahedral





VSEPR charts
Page 326-327 Table 13.1

Use the Lewis structure to determine the geometry of the molecule

e^- arrangement establishes the bond angles

Molecule takes the shape of that portion of the e^- arrangement

Charts look at the CENTRAL atom for all data!

Think REGIONS OF ELECTRON DENSITY rather than bonds (for instance, a double bond would only be 1 region)

Know these five shapes!

TABLE 12.4

Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

Case	Number of Electron Pairs	Bonds	Electron Pair Arrangement	Partial Lewis Structure	Molecular Structure	Example
1	2	2	180° Linear	A — B — A	Linear	F — Be — F BeF_2
2	3	3	120° Trigonal planar (triangular)	A, B, A	Trigonal planar (triangular)	F, B, F BF_3
3	4	4	109.5° Tetrahedral	A — B — A	Tetrahedral	H, C, H CH_4
4	4	3	109.5° Tetrahedral	A — B — A	Trigonal pyramid	H, N, H NH_3
5	4	2	109.5° Tetrahedral	A — B — A	Bent or V-shaped	H, O, H H_2O

Other VSEPR charts

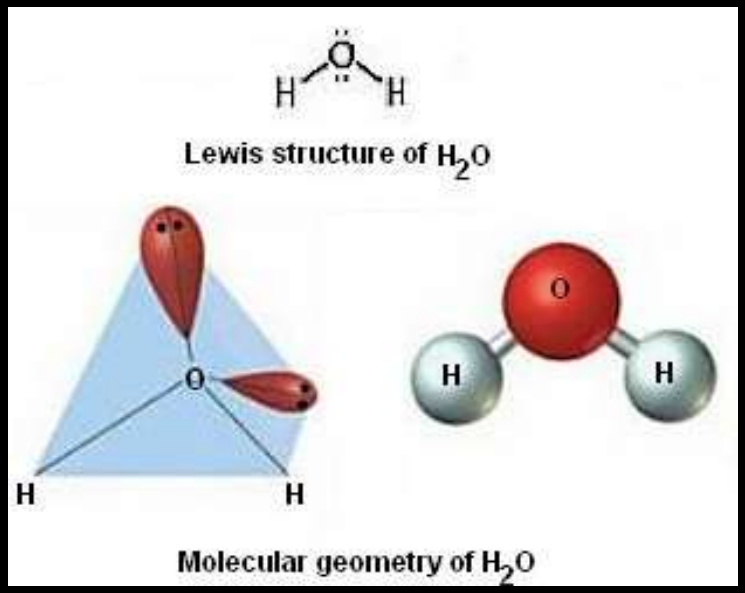
No. of Electron Pairs	Electron Pair Geometry (Bond Angle)	No. of Pendant Atoms	Molecular Geometry	Example Formula	Example Image
2	linear (180°)	2	linear	BeH_2	
3	trigonal planar (120°)	3	trigonal planar	CO_3^{2-}	
		2	bent	NO_2^-	
4	tetrahedral (109.5°)	4	tetrahedral	CH_4	
	(107 deg)	3	trigonal pyramidal	NH_3	
	(104.5 deg)	2	bent	H_2O	
5	trigonal bipyramidal (90°, 120°)	5	trigonal bipyramidal	PCl_5	
		4	see-saw	SF_4	
		3	T-shaped	BrF_3	
		2	linear	ICl_2^-	
6	octahedral (90°)	6	octahedral	SF_6	
		5	square pyramidal	BrF_5	
		4	square planar	ICl_4^-	

Table 8.6 Arrangements of Electron Pairs Around an Atom Yielding Minimum Repulsion

Regions of Electron Density	Arrangement of Electron Pairs		Example
2	Linear		
3	Trigonal planar		
4	Tetrahedral		
5	Trigonal bipyramidal		
6	Octahedral		

Structure Determination by VSEPR

Water, H_2O

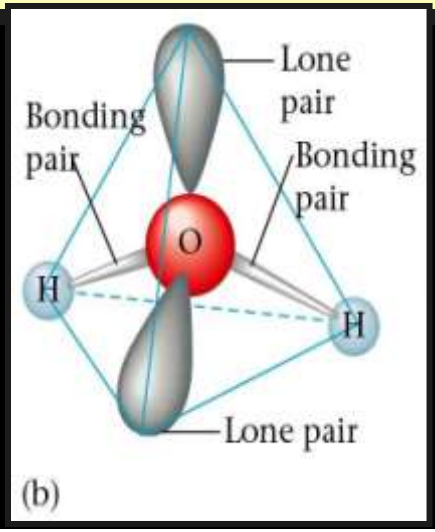


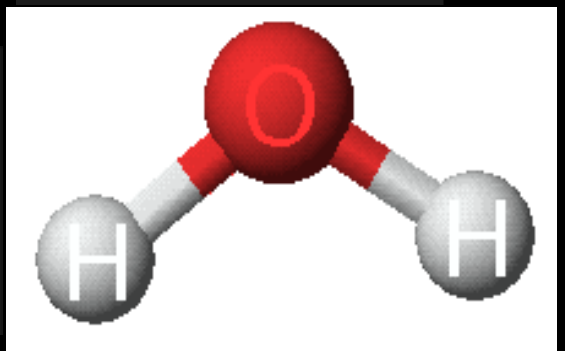
2 bond pairs

2 lone pairs

The molecular geometry is BENT.

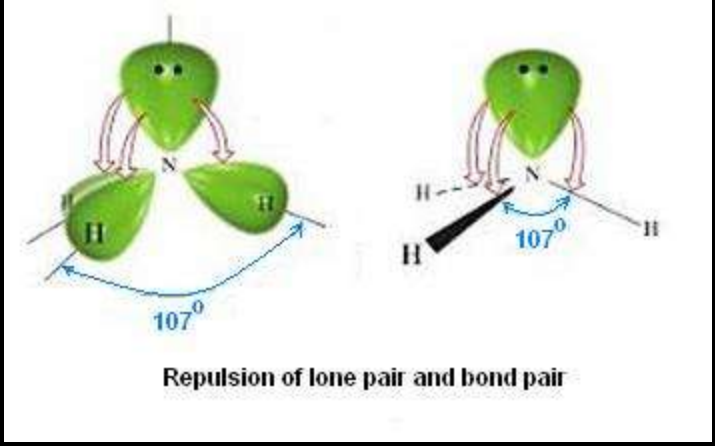
The electron pair geometry is TETRAHEDRAL





Structure Determination by VSEPR

Ammonia, NH_3
The electron pair geometry is tetrahedral.



The MOLECULAR GEOMETRY — the positions of the atoms — is TRIGONAL PYRAMID.